FIGURE A3.3 A very fast chemical reaction produces the gas that inflates the bag.

An automobile air bag is a good example of a commercial use of an explosion in a safety device (Figure A3.3). A solid compound called sodium azide ($NaN_{3(s)}$) is heated and ignited by an electrical signal when a front-end collision occurs. The air bag inflates in about 25 milliseconds, just in time to cushion the body of a person 30 to 50 milliseconds after the collision. In this case, nitrogen gas is produced when a solid suddenly decomposes.

Reactions That Form Solids

A common household application that forms a solid is a popular glue known as five-minute epoxy. The product package contains two chemicals: a resin and a hardener. When you are ready to use the glue, you mix the two chemicals together. You can work with the product for about five minutes before it begins to harden. It fully hardens in one hour. Each of the molecules in the resin has several places where it can join with the molecules in the hardener. A web of links forms, causing the whole mixture to form one large molecule. The result is a solid that will hold its shape and not melt. The bonds are strong and will bind many materials that are otherwise difficult to glue together, such as plastic to steel.

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The following symbols are used to indicate states in chemical formulas:

(s) solid
(l) liquid
(g) gas
(aq) aqueous (dissolved in water)

Showing States in Chemical Formulas

You have probably noticed that the states of each substance in a chemical reaction are often provided. These are included to give as much information as possible about a reaction. You can assume that the state subscript refers to the substance's state at room temperature (25°C), unless other information is provided. For example, at 25°C, carbon dioxide is a gas and is written as $CO_{2(g)}$, but "dry ice"—frozen carbon dioxide—is solid carbon dioxide, and should be written as $CO_{2(s)}$. Table salt is a solid ($NaCl_{(s)}$), but when dissolved in water, it is aqueous ($NaCl_{(aq)}$). Here are some guidelines for the states of substances at room temperature.

Elements

- Metals are solid, except mercury, which is a liquid.
- Most of the diatomic elements are gases: $H_{2(g)}$, $N_{2(g)}$, $O_{2(g)}$, $F_{2(g)}$, and $Cl_{2(g)}$. Bromine is a liquid, and iodine is a solid: $Br_{2(l)}$ and $I_{2(s)}$.
- Sulfur, phosphorus, and carbon are solids.

Compounds

- All ionic compounds are solid at room temperature.
- An ionic compound that is very soluble is shown as aqueous when it is dissolved in water. An ionic compound that is slightly soluble is usually shown as a solid, even when it's in water.
- Molecular compounds are very difficult to predict. The smaller the molecules are, the more they tend to be gases. The larger they are, the more they tend to be liquids and then solids. For example, $CH_{4(g)}$ is a gas (natural gas), $C_6H_{14(l)}$ is a liquid (component of gasoline), and $C_{18}H_{38(s)}$ is a solid (bees wax).

Energy Changes

Energy flow is an essential part of any chemical reaction. In some reactions, energy is released. Earlier in this section, you saw examples of how the rapid release of energy can produce explosions. In other chemical reactions, energy is absorbed.

Exothermic Reactions

Exothermic reactions release energy, usually in the form of heat, light, or electricity. The lead-acid storage battery in Figure A3.4 is an important commercial product. It discharges during ignition and charges while the vehicle is running. While it is discharging, an exothermic reaction occurs. This type of battery has been used in almost every car and truck for over 100 years. It can withstand considerable shaking and is one of the few batteries that can be recharged repeatedly—several thousand times! The battery contains sulfuric acid, which works as an electrolyte to carry an electric current between the solid chemicals inside the battery. The solids are lead and lead(IV) oxide. They react in reversible chemical reactions inside the battery. When the vehicle is started, the battery releases energy in the form of electricity. As the vehicle is running, the battery recharges.

Each discharge and recharge of the battery wears it out as the chemicals gradually break down. Eventually, the battery must be discarded, but the lead, plastic parts, and even the sulfuric acid are recycled. Over 90% of all lead batteries are recycled, making it one of the most recycled products.

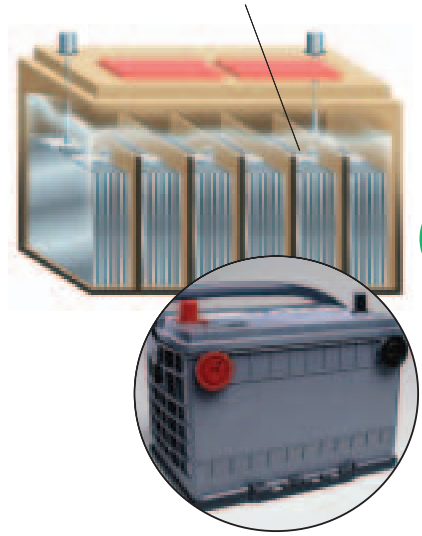

FIGURE A3.4 Almost every car and truck produced in the last century has used the chemical reaction in a lead-acid battery like this one.

Another important exothermic reaction is the combustion of fossil fuels: coal, oil, and natural gas. **Combustion** is a chemical reaction that occurs when oxygen reacts rapidly with a substance to form a new substance and gives off energy. This is called “burning.” For example, coal is used to produce electricity. The heat released by coal combustion is used to make steam, which drives turbines that produce electricity. This process also produces carbon dioxide, which is a greenhouse gas that contributes to climate change.

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} + \text{energy}$$

coal + oxygen ⟶ carbon dioxide + energy

Automobiles use gasoline instead of coal for combustion. Gasoline is a mixture of compounds, including hexane ($C_6H_{14(l)}$). The combustion of hexane is shown in the equation below.

$$2\,C_6H_{14(l)} + 19\,O_{2(g)} \longrightarrow 12\,CO_{2(g)} + 14\,H_2O_{(g)} + \text{energy}$$

hexane + oxygen ⟶ carbon dioxide + water + energy

Endothermic Reactions

Endothermic reactions absorb energy. For example, a cold pack, like the one in Figure A3.5, contains chemicals that absorb energy directly from the environment. When you squeeze the package, you break a container inside the pack that keeps the chemicals separate from each other. When the container breaks, the chemicals mix and react. As they react, they absorb energy, and the whole mixture cools down.

FIGURE A3.5 An athelete cools an injury using a cold pack. The chemicals in the cold pack are reacting in an endothermic reaction.

Biochemical Reactions

The reactions we have looked at so far are important for our society. For example, combustion of fossil fuels produces energy for transportation. The next group of reactions we will look at are biochemical reactions. They are essential to life itself.

Biochemical reactions may be endothermic or exothermic. They happen at an organism's internal temperature, and they are almost always helped by enzymes (biological catalysts). Catalysts are chemicals that speed up a reaction but are not used up by it. One of the most important chemical reactions on Earth is photosynthesis. Life on Earth depends on photosynthesis. In earlier grades, you learned that photosynthesis is a complex process that captures the Sun's light and allows a plant to use this energy to make sugar molecules. It is an endothermic reaction, as shown below. In an endothermic reaction, energy appears on the left side of the equation.

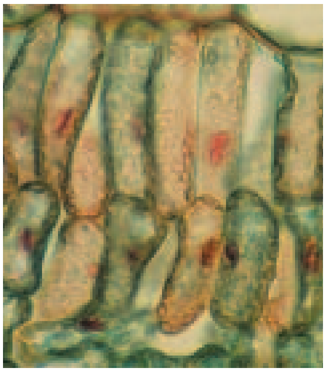

FIGURE A3.6 The chemical reaction of photosynthesis is essential to life on Earth. It takes place in chloroplasts in the cells of green leaves and stems.

$$\text{energy} + 6\,CO_{2(g)} + 6\,H_2O_{(l)} \longrightarrow C_6H_{12}O_{6(aq)} + 6\,O_{2(g)}$$

$$\text{energy} + \text{carbon dioxide} + \text{water} \longrightarrow \text{glucose} + \text{oxygen}$$

Almost all food on Earth begins with this reaction in the leaves of plants (Figure A3.6). The oxygen is a byproduct of the reaction. Most of the oxygen in our atmosphere comes from photosynthesis.

Cellular respiration is the reverse of photosynthesis. Both plants and animals use cellular respiration to release energy that is then used to drive all the chemical reactions in their tissues and organs. The equation for this exothermic reaction is shown below. This is an exothermic reaction so energy appears on the right side of the equation.

$$C_6H_{12}O_{6(aq)} + 6\,O_{2(g)} \longrightarrow \text{energy} + 6\,CO_{2(g)} + 6\,H_2O_{(l)}$$

$$\text{glucose} + \text{oxygen} \longrightarrow \text{energy} + \text{carbon dioxide} + \text{water}$$

Skill Practice — Making Inferences

In Activity A9, you will be making inferences. An inference is a conclusion made by analyzing facts. When you draw conclusions about the observations you make in a scientific investigation, you are making inferences. An inference is a logical analysis of facts, so it can always be justified by those facts.

For example, an advertisement states that 20 out of 25 people prefer Brand A cola over Brand B. Can you infer that 80 percent of all people prefer Brand A cola? No, because you do not know how many people were interviewed or how these people were chosen. Were they chosen at random, or were they all regular buyers of Brand A cola? Without this information, there is not enough data to make an inference.

For each of the following situations, write an inference based on the given data. If there isn't enough data to justify an inference, write a sentence to explain why.

1. The juice stored in the back of the bottom shelf of the refrigerator is frozen. What can you infer about the temperature in the refrigerator?
2. Your cake comes out of the oven looking more like a pancake than a light, fluffy cake. What can you infer about the length of the baking time?
3. Eight in ten dentists recommend Brand X toothpaste for reducing cavities. What can you infer about this toothpaste?

Activity A9

Inquiry Lab

Student Reference 2

Required Skills
- Initiating and Planning
- Performing and Recording
- Analyzing and Interpreting
- Communication and Teamwork

Mass Change in Chemical Reactions

The Question

Does the mass of a system change during a chemical reaction?

The Hypothesis

Read over the procedure and write a hypothesis about what happens to the mass of a system during a chemical reaction.

Variables

Read over the procedure, and identify the manipulated, responding, and controlled variables.

Materials and Equipment

- sodium carbonate solution
- calcium chloride solution
- dilute hydrochloric acid
- sodium hydrogencarbonate powder
- Erlenmeyer flask with tightly fitting stopper
- small test tube
- 50-mL beaker
- scupula
- balance
- 25-mL graduated cylinder

Procedure

Part 1: Reaction in a Sealed Apparatus

1. Test that the equipment fits together by assembling it as follows. Place a small test tube inside the Erlenmeyer flask and seal the flask with the stopper. Verify that the stopper fits snugly and does not fall out when the apparatus is inverted. Take the test tube out again.
2. Use the graduated cylinder to pour 25 mL of the calcium chloride solution into the Erlenmeyer flask. Wipe the outside of the flask to make sure it is dry. Set it aside.
3. Fill the test tube about 80% full with the sodium carbonate solution. Wipe the outside of the test tube to make sure it is dry.
4. Carefully place the filled test tube into the Erlenmeyer flask. Carefully seal the flask with the stopper, ensuring that no mixing of the solutions occurs. Measure the mass of the sealed assembly and record it.
5. Tip the assembly so that the two liquids can mix. Make sure that the stopper does not come off or allow liquid to leak. Observe the reaction and record your observations.
6. Predict whether the mass of the assembly has decreased, stayed the same, or increased. Record your prediction.
7. Measure the mass of the assembly again and record it.
8. Follow your teacher's instructions for disposing of the solutions.

Part 2: Reaction in an Open Apparatus

9. Measure 5 g of sodium hydrogencarbonate powder into a dry 50-mL beaker.
10. Fill a small test tube about 80% full with dilute hydrochloric acid. Wipe the outside of the test tube to make sure it is dry. Place the test tube in the beaker. Carefully measure and record the mass of the whole assembly without spilling any liquid from the test tube.
11. Empty the contents of the test tube into the flask and record your observations.
12. Predict whether the mass of the assembly has decreased, stayed the same, or increased. Record your prediction.
13. Find the mass of the assembly again and record it.
14. Follow your teacher's instructions for disposing of the solutions.

Analyzing and Interpreting

1. What evidence is there that a chemical reaction took place in each case?
2. How did the mass before mixing compare with the mass after mixing in each case? How did the results compare with your predictions of how the mass would change?
3. Usually the two reactions give different results with regard to mass changes. Taking into account the observations that you made, explain why you would expect these results to be different.

Forming Conclusions

4. Use your observations and the data collected to answer the question posed at the beginning of the activity.

Characteristics of Chemical Reactions

At the beginning of this unit, you learned how to recognize chemical reactions. As you work through the unit, you are learning more about them. The following points summarize the characteristics of all chemical reactions.

- All reactions involve the production of new substances with their own characteristic properties. These properties include: state at room temperature, melting point, colour, and density.
- All reactions involve the flow of energy. This may be detected by a change in temperature during the reaction. Endothermic reactions absorb energy from the environment. Exothermic reactions release energy to the environment.
- When new substances form in chemical reactions, sometimes changes of state can be observed, for example, formation of a gas (bubbles) or a solid (precipitate).
- All chemical reactions are consistent with the law of conservation of mass.

Conservation of Mass

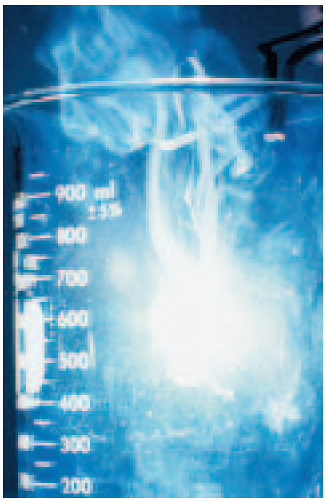

FIGURE A3.7 Magnesium burning in oxygen

In 1789, the French chemist Antoine Lavoisier came to an important conclusion based on the results of many experiments he had done. He concluded that, regardless of the type of chemical reaction, the total mass of the reaction system never changes. The "reaction system" includes both the reactants and the products. Here is another way of saying this: When a system of chemicals reacts completely, the total mass of all of the reactants equals the total mass of the products. A shorter way of saying this is that conservation of mass occurs in all chemical reactions ("to conserve" means to stay constant).

Very quickly, people recognized that Lavoisier's conclusion has many applications. For example, suppose 23.0 g of magnesium metal is burned in pure oxygen, as shown in Figure A3.7. A white powder forms. It is magnesium oxide. When all of the white powder is carefully collected and placed on a scale, its mass is measured as 39.0 g. That is, we have 16.0 g more white powder after the reaction than we had magnesium metal to start: $39.0\text{ g} - 23.0\text{ g} = 16.0\text{ g}$.

Does this mean that Lavoisier's conclusion is wrong? No, it isn't wrong. In fact, it's providing us with information. Using Lavoisier's conclusion, we can determine that 16.0 g of oxygen reacted with the 23.0 g of magnesium metal. What if we had not yet discovered oxygen? The difference in mass indicates that there must be a reactant we can't see—some new form of matter. The difference in mass between the magnesium and white powder product also gives us the mass of this unseen compound.

No exceptions to Lavoisier's proposal about the conservation of mass during a chemical reaction are known. Not all situations are testable, of course. A forest fire appears to cause a lot of mass to disappear, as the trees and other organic matter burn up. However, the combustion of wood has been carefully tested in the lab. These tests have shown that the mass of the reactants equals the mass of the products. The mass of the gases produced plus the mass of the ash equals the mass of the wood and the oxygen that were present before combustion began.

In science, observations that have wide application and appear to have no exceptions are given a special status—they are called laws. Antoine Lavoisier's conclusion is called the **law of conservation of mass**. Lavoisier did not know of the existence of atoms because there was very little evidence for them at the time. Recall from section A1.0 that people were still developing different models of matter. However, we do know about atoms today. Using the law of conservation of mass, we can deduce that the total number of atoms present before a reaction is equal to the total number of atoms present after a reaction. This is important in writing chemical equations, as you will be doing later in this section.

reSEARCH

Investigate the proportion of greenhouse gases in Alberta that are produced from forest fires, compared with other sources. Begin your search at www.pearsoned.ca/school/science10

A3.1 Check and Reflect

Knowledge

1. List five industrial or commercial processes that use chemical reactions.

2. Explain the difference between the meanings of the state symbols *(l)* and *(aq)*.

3. How are the combustion of coal and the combustion of hexane (in gasoline) similar in terms of the products of these reactions?

4. State the law of conservation of mass.

5. What is the difference between an exothermic and an endothermic reaction?

6. Explain how photosynthesis and cellular respiration are related in terms of the chemicals involved and energy.

Applications

7. Many reactions done in laboratories at high temperatures occur in living organisms at much lower temperatures. What do living systems have that allow these reactions to occur at the lower temperatures?

8. List three features of a chemical reaction that would make it suitable for propelling a rocket.

9. What are the environmental impacts of the combustion of fossil fuels?

10. A 20.2-g sample of carbon dioxide contains 5.50 g of carbon. What mass of oxygen is present in the sample?

11. A 100.0-g sample of sugar contains carbon, hydrogen, and oxygen. The sample contains 40.0 g of carbon and 53.3 g of oxygen. What mass of hydrogen is in the sample?

Extensions

12. Iron and sulfur react when heated together. If 50.0 g of iron and 100.0 g of sulfur are mixed together and heated, a product with a mass of 107.4 g is produced. All of the iron reacted, and excess sulfur burned off during the process.

a) What mass of sulfur combined chemically with the iron?

b) What mass of sulfur was burned off?

13. In principle, all chemical reactions are reversible; that is, under the right circumstances, the products of a reaction can be used to make the reactants again. Photosynthesis and cellular respiration are related to each other this way. Use this idea to explain why, for every exothermic reaction in the universe, there is a corresponding endothermic one.

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The first noble gas compound ever synthesized was produced in 1962 at the University of British Columbia. Xenon reacted with fluorine gas to form the yellow-orange crystalline solid xenon difluoride. The chemical reaction is written as follows:

$$Xe_{(g)} + F_{2(g)} \longrightarrow XeF_{2(s)}$$

FIGURE A3.8 Evidence for the reaction between nitric acid and copper comes from the formation of bubbles, a yellow-brown gas, and a green or blue solution.

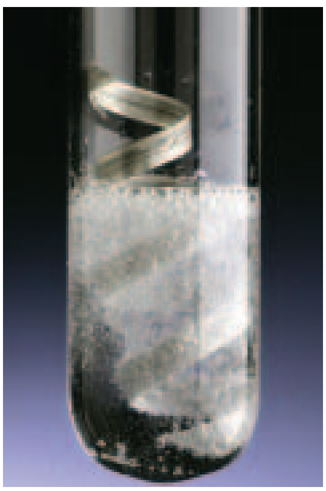

FIGURE A3.9 The reaction of magnesium with hydrochloric acid results in an increase in temperature.

A 3.2 Writing Chemical Equations

Recall that a chemical reaction is a process in which one or more substances are transformed into new substances, each with its own properties. The process also involves energy being released or absorbed. In other words, a chemical reaction is a process involving chemical change. Chemists record such a process in a **chemical equation** that uses chemical symbols and formulas. This equation is a shorthand way of showing what happens during a reaction. Other special symbols are also used. You have seen many chemical equations in this unit. Now you will learn how to write them yourself.

Symbolizing Chemical Change

To write a chemical equation, you need to know what substances react (the reactants) and what new substances form (the products). This requires the following:

- careful observations
- knowledge of what substances are present at the start of the reaction
- the ability to analyze the materials produced by the reaction

The first step in writing a chemical equation is to recognize that a chemical change has actually occurred. To do this, look for changes in properties and changes in energy.

Consider the reaction between copper and nitric acid in Figure A3.8. The photo shows two substances that were not present in the unmixed reactants: a brown gas and a green solution. The presence of bubbles indicates that a gas is being produced. Closer inspection would reveal that the pennies are corroding, and the mixture has become very hot.

Chemical change is often accompanied by visible events, such as the production of gas (bubbling), the release of heat (increased temperature), a change in colour, or the appearance of a substance that is only slightly soluble (cloudiness).

Writing Word Equations

Consider the chemical reaction of a piece of magnesium metal with hydrochloric acid (Figure A3.9). The corrosion of the magnesium and fizzing of the liquid are evidence that a reaction is taking place. The temperature is elevated above room temperature, showing that this is an exothermic reaction. We can describe the reaction with the following sentence:

> Solid **magnesium** metal reacts with aqueous **hydrochloric acid** to produce aqueous **magnesium chloride** and **hydrogen** gas.

The starting materials (in blue) are the reactants. The substances made during the reaction (in red) are the products. The physical states (e.g., solid, gas) of each substance in the reaction are also mentioned. "Aqueous" means that the magnesium chloride is dissolved in water. The word equation for this reaction is:

$$\text{magnesium} + \text{hydrochloric acid} \longrightarrow \text{magnesium chloride} + \text{hydrogen gas}$$

A plus sign (+) groups the reactants together. It does not matter what order the reactants are written. Hydrochloric acid could have been written first. The arrow ($\longrightarrow$) separates the reactants from the products and is read "produces." The products are also joined by a "+" sign.

The use of word equations opens the door to easily describing countless examples of chemical change. Figure A3.10 shows another example of a chemical reaction. An iron nail is placed in a solution of copper(II) chloride. The equation for the reaction is:

iron + copper(II) chloride $\longrightarrow$ iron(II) chloride + copper

FIGURE A3.10 An iron nail in a solution of copper(II) chloride

You may be wondering how it is possible to know what the products of a chemical reaction are. With experience, you will often be able to predict what will happen in a given reaction just by looking at the names of the reactants. This prediction is easier if you use formula equations. Formula equations also help you describe chemical changes more precisely.

Writing Balanced Formula Equations

The bus in Figure A3.11 runs on electricity produced by a fuel cell. The electricity comes from a reaction between hydrogen and oxygen to form water. The reaction can be described with the following word equation:

hydrogen + oxygen $\longrightarrow$ water

FIGURE A3.11 Hydrogen and oxygen gas combine chemically in a fuel cell. The reaction in the fuel cell produces electrical energy to operate the bus.

A **formula equation** uses the chemical formulas of reactants and products in a chemical equation to represent a chemical reaction. Recall that both oxygen and hydrogen gases exist as diatomic molecules. Their formulas are $O_{2(g)}$ and $H_{2(g)}$. The formula equation for this reaction is:

$$H_{2(g)} + O_{2(g)} \longrightarrow H_2O_{(l)}$$

This equation is called a **skeleton equation**. It shows the identities of the substances involved in the reaction, and which elements are present. A skeleton equation is of limited value, because it does not show the correct proportions of the reactants and the products.

How do we know what these correct proportions are? The law of conservation of mass gives us this information. Recall that the law states that the total mass of the reactants in a reaction equals the total mass of the products. The mass of all the components is represented by the number of atoms of each element in the reactants and products. To show the correct proportions, we need to write a balanced equation. The balanced equation for the formation of water is:

$$2\ H_{2(g)} + O_{2(g)} \longrightarrow 2\ H_2O_{(l)}$$

This equation is also shown in Figure A3.12

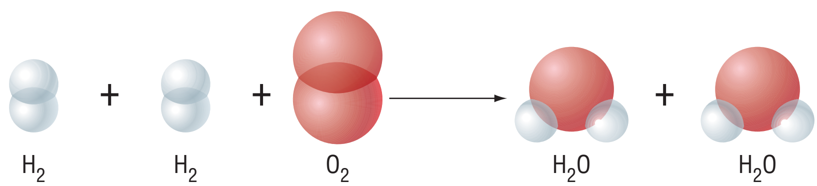

FIGURE A3.12 Molecules of hydrogen and oxygen react to form water.

FIGURE A3.13 Methane combustion provides warmth in a natural gas fireplace or furnace.

Notice that integers have been placed in front of the formulas for the hydrogen and the water molecules. These are called coefficients. In this case, the coefficient 2 is used to ensure that the number of hydrogen and oxygen atoms in the reactants equals the number of hydrogen and oxygen atoms in the products. Oxygen does not need a coefficient in front of it because only one molecule ($O_{2(g)}$) is needed on the reactant side of the equation to balance the two oxygen atoms in the products.

An equation is properly balanced if the number of each type of atom on the reactants side of an equation is equal to the number of each type of atom on the products side. Consider another reaction involving oxygen, but this time, oxygen is reacting with methane gas (Figure A3.13). The word equation for this reaction is:

oxygen + methane $\longrightarrow$ carbon dioxide + water

To write the skeleton equation, you need to know that oxygen is diatomic ($O_{2(g)}$) and that the formula for methane is $CH_{4(g)}$. You can find formulas for some compounds in Table B of Student Reference 12. The skeleton equation for this reaction is:

$$O_{2(g)} + CH_{4(g)} \longrightarrow CO_{2(g)} + H_2O_{(g)}$$

The balanced equation is:

$$2\ O_{2(g)} + CH_{4(g)} \longrightarrow CO_{2(g)} + 2\ H_2O_{(g)}$$

Note that you balance an equation by making sure that the number of atoms of each element left of the arrow is equal to the number of atoms of each element right of the arrow. Note also that you *cannot* change the formulas of any of the substances, so you cannot balance the hydrogen atoms by removing two from the subscript beside the H in $CH_{4(g)}$. Never change a subscript to balance an equation. Instead, add coefficients. A balanced equation contains coefficients that consist of the lowest whole-number ratios of the substances involved in the reaction.

Table A3.1 shows how atoms are counted.

TABLE A3.1 Number of Atoms in the Equation for the Combustion of Methane

Combustion of Methane: $2\ O_{2(g)} + CH_{4(g)} \longrightarrow CO_{2(g)} + 2\ H_2O_{(g)}$		
Type of Atom	**Number of Reactant Atoms (coefficient × no. of atoms)**	**Number of Product Atoms (coefficient × no. of atoms)**
O	$2 \times 2 = 4$	$1 \times 2 + 2 \times 1 = 4$
C	1	1
H	4	$2 \times 2 = 4$

Example Problem A3.1

Aqueous iron(II) nitrate reacts with aqueous sodium phosphate. The products are aqueous sodium nitrate and solid iron(II) phosphate. Iron(II) phosphate is used as a fertilizer to prevent iron deficiency in trees. This condition prevents trees from making chlorophyll. Write the balanced equation for this reaction, and include symbols showing the states.

1. Write the word equation for the reaction.

iron(II) nitrate + sodium phosphate $\longrightarrow$ sodium nitrate + iron(II) phosphate

2. Write the skeleton equation for the reaction.

$$Fe(NO_3)_{2(aq)} + Na_3PO_{4(aq)} \longrightarrow NaNO_{3(aq)} + Fe_3(PO_4)_{2(s)}$$

3. Count the number of each type of atom or polyatomic ion in the reactants and the products. It may be helpful to use a table like Table A3.2.

TABLE A3.2 Number of Atoms and Polyatomic Ions in the Reaction

Reaction of Iron(II) Nitrate with Sodium Phosphate		
Atom or Polyatomic Ion	**Reactants**	**Products**
Fe	1	3
Na	3	1
NO_3^-	2	1
PO_4^{3+}	1	2

4. Remember that you cannot change the formulas of any of the substances. Treat polyatomic ions as single units. Add coefficients:

- Start with the first element on the left, Fe. It is not balanced. Place a 3 in front of $Fe(NO_3)_2$ in the reactants to balance it with the Fe in the $Fe_3(PO_4)_2$ product.
- NO_3 appears on both sides of the equation. Place a 6 in front of $NaNO_3$ to balance the NO_3 on both sides. This 6 means there are now 6 Na atoms on the product side.
- Place a 2 in front of the Na_3PO_4 to balance the number of Na atoms. The PO_4 polyatomic ions are now also balanced.

5. The result is the balanced equation:

$$3\ Fe(NO_3)_{2(aq)} + 2\ Na_3PO_{4(aq)} \longrightarrow 6\ NaNO_{3(aq)} + Fe_3(PO_4)_{2(s)}$$

Practice Problem

1. Balance the following equations:

a) $N_{2(g)} + H_{2(g)} \longrightarrow NH_{3(g)}$

b) $CaC_{2(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2(s)} + C_2H_{2(g)}$

c) $SiCl_{4(s)} + H_2O_{(l)} \longrightarrow SiO_{2(s)} + HCl_{(aq)}$

d) $H_3PO_{4(aq)} + CaSO_{4(s)} \longrightarrow Ca_3(PO_4)_{2(s)} + H_2SO_{4(aq)}$

reSEARCH

Catalysts are chemicals that can speed up a chemical reaction, but are not changed by it. A catalyst is present at the end of a reaction in the same amount as at the start of the reaction. Find out how a catalyst is represented in a balanced chemical reaction. Begin your search at www.pearsoned.ca/school/science10

A3.2 Check and Reflect

Knowledge

1. What is a chemical equation?
2. List four observations that could indicate that a chemical reaction is taking place.
3. List four pieces of information given by a balanced formula equation.
4. What information about chemical reactions is not given by the chemical equation?
5. Which law of nature are we using when we balance a chemical equation?
6. Use the following equation to explain the following terms: reactants, products, state symbols, formulas, coefficients.

   $Zn_{(s)} + 2\ HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2(g)}$

Applications

7. Balance the following equations:
   a) $Al_{(s)} + F_{2(g)} \longrightarrow AlF_{3(s)}$
   b) $K_{(s)} + O_{2(g)} \longrightarrow K_2O_{(s)}$
   c) $C_6H_{12}O_{6(s)} + O_{2(g)} \longrightarrow CO_{2(g)} + H_2O_{(g)}$
   d) $H_2SO_{4(aq)} + NaOH_{(s)} \longrightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$
   e) $Mg(CH_3COO)_{2(aq)} + AgNO_{3(aq)} \longrightarrow Mg(NO_3)_{2(aq)} + AgCH_3COO_{(s)}$
   f) $H_2O_{2(aq)} \longrightarrow O_{2(g)} + H_2O_{(l)}$
8. For each of the following, write skeleton formula equations, and then balance them.
   a) methane + oxygen ⟶ carbon dioxide + water vapour
   b) sodium chloride ⟶ sodium + chlorine
   c) calcium nitrate + sodium sulfate ⟶ sodium nitrate + calcium sulfate
   d) hydrogen + carbon monoxide ⟶ carbon + water
   e) sodium + water ⟶ sodium hydroxide + hydrogen
   f) calcium carbonate + sulfur dioxide + oxygen ⟶ calcium sulfate + carbon dioxide
   g) sulfur + oxygen ⟶ sulfur dioxide
   h) calcium phosphate + sulfuric acid ⟶ phosphoric acid + calcium sulfate
   i) potassium chlorate ⟶ potassium chloride + oxygen
9. Write the following equations as balanced formula equations.
   a) Solid calcium metal is placed in a solution of hydrochloric acid, producing aqueous calcium chloride and hydrogen gas.
   b) Solid magnesium nitride is placed in water and stirred. This produces aqueous magnesium hydroxide and ammonia gas.
   c) Aqueous sulfuric acid reacts with solid sodium hydroxide to produce aqueous sodium sulfate and liquid water.
   d) Gaseous nitrogen dioxide reacts with itself to produce gaseous dinitrogen tetraoxide.
   e) Aqueous copper(II) chloride mixes with aqueous sodium hydroxide to produce solid copper(II) hydroxide plus aqueous sodium chloride.

Extension

10. Create a drawing to illustrate the law of conservation of mass by showing how the atoms rearrange during the decomposition of hydrogen peroxide into oxygen and water.

A 3.3 Five Common Types of Chemical Reactions

How many different chemical reactions are possible? There are millions of chemical compounds, and each one can undergo many different kinds of chemical change. It would be impossible to learn all these reactions. Fortunately, chemistry is rich in patterns. Chemists looked at many different reactions and found that some had common characteristics. From the vast array of reactions, a few simple types have emerged. They allow us to predict the outcome of many chemical reactions just by examining the reactants.

Five common types of reactions are: formation, decomposition, hydrocarbon combustion, single replacement, and double replacement. You will explore them in more detail in the rest of section A3.3.

Formation Reactions

In the simplest type of **formation reaction**, two elements combine to form a compound. This type of reaction is also known as a **synthesis reaction**. A general statement for a formation reaction is:

element + element $\longrightarrow$ compound

OR

$A + B \longrightarrow AB$

When sulfur burns in air, the reaction produces a poisonous gas called sulfur dioxide (Figure A3.14). The equation for this reaction is:

Word equation: sulfur + oxygen $\longrightarrow$ sulfur dioxide

Skeleton equation: $S_{8(s)} + O_{2(g)} \longrightarrow SO_{2(g)}$

Sulfur dioxide can combine with water in the air to form acid rain. That is why as much sulfur as possible is removed from gasoline during production.

Note that in this section, we will often use only skeleton equations and not balanced equations. This allows us to focus on the reaction types.

Another formation reaction involving only non-metals is the synthesis of ammonia from its elements. The formula of ammonia is NH_3, so the reactants must be the elements hydrogen and nitrogen. Both of these elements form diatomic molecules.

Word equation: hydrogen + nitrogen $\longrightarrow$ ammonia

Skeleton equation: $H_{2(g)} + N_{2(g)} \longrightarrow NH_{3(g)}$

Because ammonia is used in the production of explosives and fertilizers, its formation is an important commercial reaction. The reaction is carried out at high temperatures and pressure, in the presence of a catalyst. Note that you cannot always predict the formula of molecular compounds that form because more than one combination of elements is possible. For example, carbon and oxygen can react to form $CO_{(g)}$ or $CO_{2(g)}$.

*info*BIT

The car of the future may operate on hydrogen instead of gasoline. Burning gasoline produces carbon dioxide, a known greenhouse gas. A hydrogen fuel cell produces only water. A hydrogen fuel cell is a device that uses a formation reaction involving hydrogen gas and oxygen gas to produce electrical energy—and water. The electrical energy can be used to drive an electric motor.

FIGURE A3.14 Sulfur burning in air is an example of a formation reaction.

FIGURE A3.15 Burning sodium metal in chlorine gas produces solid sodium chloride.

Many formation reactions occur between metals and non-metals. Table salt can be produced from two highly reactive substances: chlorine gas and sodium metal. Commercial table salt is not prepared this way, since it can easily be mined, but the reaction is possible (Figure A3.15). The equation for the reaction is:

Word equation: sodium + chlorine $\longrightarrow$ sodium chloride

Skeleton equation: $Na_{(s)} + Cl_{2(g)} \longrightarrow NaCl_{(s)}$

Example Problem A3.2

Special high quality grades of magnesium oxide are used in cosmetics, antacids, Sun blocks, and ointments. Write the balanced equation for the formation of solid magnesium oxide from its elements.

1. Write the word equation: From the name of the product compound, you can tell that the elements it is composed of are magnesium, which is a metal, and oxygen, which is a diatomic gas.

   Word equation: magnesium + oxygen $\longrightarrow$ magnesium oxide

2. Write the skeleton equation: Metals are indicated in an equation by their element symbol, and oxygen is a diatomic element. Metals other than mercury are solids at room temperature. Magnesium oxide is an ionic compound. All ionic compounds are solids at room temperature.

   Skeleton equation: $Mg_{(s)} + O_{2(g)} \longrightarrow MgO_{(s)}$

3. Balance the equation: The balanced equation for the formation of solid magnesium oxide is:

   $2\ Mg_{(s)} + O_{2(g)} \longrightarrow 2\ MgO_{(s)}$

Example Problem A3.3

Artists use iron(III) chloride solutions to etch images onto copper. Write the skeleton equation and balanced equation without state symbols for the formation of iron(III) chloride from its elements.

1. Write the word equation: From the name of the product compound, you can tell that it is composed of the elements iron, which is a metal, and chlorine, which is a diatomic gas.

   Word equation: iron + chlorine $\longrightarrow$ iron(III) chloride

2. Write the skeleton equation: Metals are always indicated in an equation by their element symbol, and chlorine is a diatomic element. Iron(III) chloride is an ionic compound.

   Skeleton equation: $Fe_{(s)} + Cl_{2(g)} \longrightarrow FeCl_{3(s)}$

3. Write the balanced equation. The balanced equation for the formation of iron(III) chloride from its elements is:

   $2\ Fe_{(s)} + 3\ Cl_{2(g)} \longrightarrow 2\ FeCl_{3(s)}$

Practice Problems

2. Write the skeleton equation and balanced equation for the formation of lithium oxide from its elements.

3. Write the skeleton equation and balanced equation for the formation of lead(IV) bromide from its elements.

More complex types of formation reactions involve compounds that react to form a single product. For example, carbon dioxide can join with water to make carbonic acid. This happens to about 10% of the carbon dioxide that is forced into soda pop in a process called carbonation, which adds the fizz to the drink (Figure A3.16). The skeleton equation for this reaction is:

$$CO_{2(g)} + H_2O_{(l)} \longrightarrow H_2CO_{3(aq)}$$

When both reactants are compounds, the product of this type of formation reaction can be very difficult to predict. When one reactant is a metal and the other is a non-metal, predict the product by writing the ionic compound that they form.

FIGURE A3.16 The bubbles in carbonated beverages are carbon dioxide gas. The carbon dioxide sometimes reacts with water in the beverage to form carbonic acid ($H_2CO_{3(aq)}$).

Example Problem A3.4

The product of this reaction is used in solution to digest wood fibre in the papermaking process. Name and give the formula of the product in the following reaction:
sodium + sulfur $\longrightarrow$

These compounds form the ionic compound sodium sulfide, formula $Na_2S_{(s)}$.

Practice Problem

4. Name and give the formula of the product in each of the following reactions:
a) calcium + nitrogen $\longrightarrow$
b) silver + oxygen $\longrightarrow$
c) aluminium + fluorine $\longrightarrow$

Skill Practice Formation Reactions

Later in this section, in Activity A10, you will be asked to write the equations for a formation reaction. Before doing that activity, you can practise identifying and predicting products and writing equations by answering the following questions.

1. Name the product in each of the following reactions:
a) potassium + iodine $\longrightarrow$
b) magnesium + phosphorus $\longrightarrow$
c) cesium + chlorine $\longrightarrow$
d) calcium + oxygen $\longrightarrow$
e) aluminium + sulfur $\longrightarrow$

2. Complete each equation below. The products are all ionic compounds.
a) $Na_{(s)} + Br_{2(l)} \longrightarrow$
b) $Mg_{(s)} + F_{2(g)} \longrightarrow$
c) $Al_{(s)} + Cl_{2(g)} \longrightarrow$
d) $K_{(s)} + N_{2(g)} \longrightarrow$
e) $Ca_{(s)} + P_{4(s)} \longrightarrow$

3. Write balanced chemical equations for the following reactions. Add state symbols. You must predict the products of the last three reactions.
a) solid lithium + oxygen gas $\longrightarrow$ solid lithium oxide
b) solid aluminium + liquid bromine $\longrightarrow$ solid aluminium bromide
c) liquid mercury + solid iodine $\longrightarrow$ solid mercury(II) iodide
d) solid sodium + chlorine gas $\longrightarrow$
e) solid magnesium + nitrogen gas $\longrightarrow$
f) solid nickel + fluorine gas $\longrightarrow$

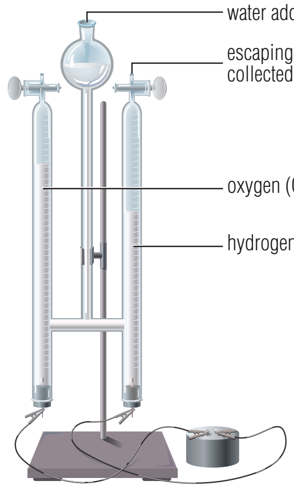

FIGURE A3.17 A Hoffman apparatus separates water into hydrogen and oxygen using electricity, in a process called electrolysis.

Decomposition Reactions

Each of the reactions described in the discussion of formation reactions above is reversible. That is, the products can be broken down to yield the reactants again in a **decomposition reaction**. Sometimes this is difficult, and requires special equipment or the input of energy, but it is possible.

A general statement for this type of reaction is:

compound $\longrightarrow$ element + element

OR

AB $\longrightarrow$ A + B

Decompostion reactions have only one reactant. One example of a decomposition reaction is the breakdown of water into hydrogen and oxygen. Figure A3.17 shows the apparatus used to do this. The balanced equation for the reaction is:

$$2\ H_2O_{(l)} \longrightarrow 2\ H_{2(g)} + O_{2(g)}$$

Example Problem A3.5

Before 1825, pure aluminium metal did not exist. In that year, a few pin-head-sized pieces of aluminium metal were produced through a reaction involving aluminium chloride. Write a balanced equation showing the decomposition of solid aluminium chloride into its elements.

1. Write the word equation: The elements in aluminium chloride are aluminium and chlorine.

Word equation: aluminium chloride $\longrightarrow$ aluminium + chlorine

2. Write the skeleton equation: Aluminium chloride is an ionic compound. All ionic compounds are solids at room temperature. Metals are indicated by their symbols. Chlorine is diatomic. All metals other than mercury are solids at room temperature, and chlorine is a gas.

Skeleton equation: $AlCl_{3(s)} \longrightarrow Al_{(s)} + Cl_{2(g)}$

3. Write the balanced equation: The balanced equation for the decomposition of solid aluminium chloride into its elements is:

$$2\ AlCl_{3(s)} \longrightarrow 2\ Al_{(s)} + 3\ Cl_{2(g)}$$

Practice Problem

5. Write balanced equations for the reactions below.

a) solid magnesium sulfide $\longrightarrow$ solid magnesium + solid sulfur
b) solid potassium iodide $\longrightarrow$ solid potassium + solid iodine
c) solid aluminium oxide $\longrightarrow$ solid aluminium + oxygen gas
d) solid nickel(II) chloride $\longrightarrow$ solid nickel + chlorine gas

Hydrocarbon Combustion

Hydrocarbons are substances that contain hydrogen and carbon. Common hydrocarbons include the main components of gasoline (which is a mixture of many liquid hydrocarbons), candle wax, and many plastics, such as polyethylene. Methane is the simplest hydrocarbon, with a formula of $CH_{4(g)}$, but there are millions of different hydrocarbon molecules.

One way to write a general formula for a hydrocarbon is C_xH_y. The x and y subscripts are whole numbers that indicate how many carbon and hydrogen atoms are in the molecule. Figure A3.18 shows how to convert the general formula of C_xH_y to the formula for a specific hydrocarbon. Butane has 4 carbon atoms and 10 hydrogen atoms, so its formula is $C_4H_{10(g)}$.

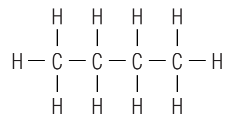

FIGURE A3.18 For butane, x = 4 and y = 10, giving the formula $C_4H_{10(g)}$.

Any reaction with oxygen that is fast and exothermic is a combustion reaction (Figure A3.19). If plenty of oxygen is available to react in hydrocarbon combustions, there will always be only two products: carbon dioxide and water vapour. The general skeleton equation for this type of reaction is:

$$C_xH_y + O_{2(g)} \longrightarrow CO_{2(g)} + H_2O_{(g)}$$

hydrocarbon + oxygen ⟶ carbon dioxide + water

Natural gas is used as a fuel for home heating and cooking. Recall that natural gas is methane, with the formula $CH_{4(g)}$. The balanced equation for the combustion of methane is:

$$CH_{4(g)} + 2\ O_{2(g)} \longrightarrow CO_{2(g)} + 2\ H_2O_{(g)}$$

FIGURE A3.19 An oxy-acetylene torch cuts through solid steel.

The following example problem uses hexane as the hydrocarbon in the combustion reaction. Notice that balancing these types of reactions often requires large numbers for the coefficients.

Example Problem A3.6

One of the hydrocarbon components of gasoline is hexane, $C_6H_{14(l)}$. Write word, skeleton, and balanced equations for the combustion of hexane.

1. Write the word equation: Combustion always means reaction with oxygen. Since hexane is a hydrocarbon, the reaction always produces carbon dioxide and water vapour.

   Word equation: hexane + oxygen ⟶ carbon dioxide + water vapour

2. Write the skeleton equation: You know from the question that hexane is a liquid (because gasoline is a liquid). The other components are gases. Except for the particular hydrocarbon, the formulas are always the same for hydrocarbon combustion.

   Skeleton equation: $C_6H_{14(l)} + O_{2(g)} \longrightarrow CO_{2(g)} + H_2O_{(g)}$

3. Write the balanced equation:

   Balanced equation: $2\ C_6H_{14(l)} + 19\ O_{2(g)} \longrightarrow 12\ CO_{2(g)} + 14\ H_2O_{(g)}$

Practice Problem

6. Complete and balance each equation.
   a) $CH_{4(g)} + O_{2(g)} \longrightarrow$
   b) $C_2H_{6(g)} + O_{2(g)} \longrightarrow$
   c) $C_3H_{8(g)} + O_{2(g)} \longrightarrow$
   d) $C_6H_{6(l)} + O_{2(g)} \longrightarrow$

Single Replacement Reactions

In a **single replacement reaction**, a reactive element reacts with an ionic compound. After the reaction, the element ends up in a compound, and one of the elements in the reactant compound ends up by itself as an element. These reactions often take place in solution. The equation for a single replacement reaction may look like this:

$$A + BC \longrightarrow B + AC$$

In one type of single replacement reaction, a metal atom trades places with a metal ion in a compound. For example, magnesium metal reacts with the compound silver nitrate as follows:

Word equation: magnesium + silver nitrate $\longrightarrow$ silver + magnesium nitrate

Any metal element on its own is present as atoms. Any metal element in a compound is present as an ion. Its ionic charge determines how the compound is written. In this example, silver is a 1+ ion in silver nitrate. In the products, silver is an element and has no charge. In the same reaction, magnesium starts out in the reactants as an element and ends up in the products as a 2+ ion. The skeleton and balanced equations for this reaction are:

Skeleton equation: $Mg_{(s)} + AgNO_{3(aq)} \longrightarrow Ag_{(s)} + Mg(NO_3)_{2(aq)}$

Balanced equation: $Mg_{(s)} + 2\ AgNO_{3(aq)} \longrightarrow 2\ Ag_{(s)} + Mg(NO_3)_{2(aq)}$

Example Problem A3.7

Copper compounds are poisonous. Sometimes, active metals such as iron or aluminium are placed in water contaminated with copper ions. These metals react with the copper ions and remove them from the water. Solid copper metal is produced, which is safe in the water. What reaction occurs when aluminium metal is placed in a solution of aqueous copper(II) chloride? Write the balanced equation.

1. Write the word equation: Aluminium metal replaces copper in the compound. Copper becomes copper metal, and aluminium forms aluminium chloride.

Word equation: aluminium + copper(II) chloride $\longrightarrow$ copper + aluminium chloride

2. Write the skeleton equation: Copper(II) chloride and aluminium chloride are both ionic compounds.

   Skeleton equation: $Al_{(s)} + CuCl_{2(aq)} \longrightarrow Cu_{(s)} + AlCl_{3(aq)}$

3. Write the balanced equation:

   Balanced equation: $2\ Al_{(s)} + 3\ CuCl_{2(aq)} \longrightarrow 3\ Cu_{(s)} + 2\ AlCl_{3(aq)}$

In another type of single replacement reaction, non-metals trade places. The equation for this reaction may look like this:

$$D + BC \longrightarrow C + BD$$

For example, chlorine reacts with the compound silver bromide to produce bromine and silver chloride. Both chlorine and bromine are

diatomic elements. Silver bromide and silver chloride are ionic compounds. The balanced equation for this reaction is:

Balanced equation: $Cl_{2(g)} + 2\ AgBr_{(s)} \longrightarrow Br_{2(l)} + 2\ AgCl_{(s)}$

Example Problem A3.8

Iron(III) bromide is an industrial catalyst. It helps to speed certain chemical reactions used in manufacturing plastics. Liquid bromine is added to a solution of aqueous iron(III) iodide, and the mixture is stirred. This produces aqueous iron(III) bromide and solid iodine. Write the word, skeleton, and balanced equations for this reaction. Include the state symbols in the balanced equation.

1. Write the word equation: Read the question carefully to find the names of the reactants and products.

Word equation: bromine + iron(III) iodide ⟶ iodine + iron(III) bromide

2. Write the skeleton equation: Bromine and iodine are both diatomic elements. The state of each substance is given in the question.

   Skeleton equation: $Br_{2(l)} + FeI_{3(aq)} \longrightarrow I_{2(s)} + FeBr_{3(aq)}$

3. Write the balanced equation: In balancing, remember not to change any subscripts; just add coefficients where needed.

   Balanced equation: $3\ Br_{2(l)} + 2\ FeI_{3(aq)} \longrightarrow 3\ I_{2(s)} + 2\ FeBr_{3(aq)}$

Practice Problems

Write the word, skeleton, and balanced equations.

7. Chlorine gas is added to a solution of aqueous nickel(III) bromide and the mixture is stirred. This produces aqueous nickel(III) chloride and liquid bromine.
8. Zinc metal is placed into a solution of silver nitrate and allowed to sit. This produces aqueous zinc nitrate and solid silver metal.

Skill Practice — Decomposition and Single Replacement Reactions

In Activity A10, you will be asked to write the equations for formation and single replacement reactions. You have already had an opportunity to practise identifying and predicting products and writing equations for formation reactions in the Skill Practice on page 93. You can practise identifying and predicting products and writing equations for decomposition and single replacement reactions by answering the following questions.

1. Name the products in each of the following reactions:
   a) magnesium phosphide ⟶
   b) sodium chloride ⟶
   c) strontium oxide ⟶
   d) zinc + iron(II) chloride ⟶
   e) aluminium + copper(II) iodide ⟶
   f) magnesium + gold(III) nitrate ⟶
2. Complete each skeleton equation.
   a) $CaO_{(s)} \longrightarrow$
   b) $NaF_{(s)} \longrightarrow$
   c) $Mg_3N_{2(s)} \longrightarrow$
   d) $Fe_{(s)} + Cu(NO_3)_{2(aq)} \longrightarrow$
   e) $Cl_{2(g)} + NaI_{(aq)} \longrightarrow$
   f) $Pb_{(s)} + AgNO_{3(aq)} \longrightarrow$
3. Write balanced chemical equations for the following reactions. You must predict the products where indicated.
   a) solid iron(III) chloride ⟶ solid iron + chlorine gas
   b) solid copper(I) oxide ⟶ solid copper + oxygen gas
   c) solid lithium bromide ⟶ ?
   d) liquid bromine + aqueous chromium(III) iodide ⟶ chromium(III) bromide + solid iodine
   e) aqueous silver nitrate + solid copper ⟶ ?
4. Write the word, skeleton, and balanced equations for the following reactions. Include the state symbols in the balanced equations.
   a) Liquid bromine is added to a solution of aqueous iron(III) iodide and the mixture is stirred. This produces aqueous iron(III) bromide and solid iodine.
   b) Magnesium metal is placed into a solution of gold(III) fluoride and allowed to sit. This produces aqueous magnesium fluoride and gold metal.

Activity A10

Inquiry Lab

Student Reference 2

Required Skills
- Initiating and Planning
- Performing and Recording
- Analyzing and Interpreting
- Communication and Teamwork

Formation, Decomposition, and Single Replacement Reactions

The Question

What chemical changes happen during formation, decomposition, and single replacement reactions?

Materials and Equipment

- anhydrous copper(II) sulfate
- 3% hydrogen peroxide solution
- manganese(IV) oxide
- 15-cm length of copper wire
- steel wool
- water
- balance
- test tube
- medicine dropper
- tongs
- Bunsen burner
- Erlenmeyer flask
- copper(II) sulfate pentahydrate
- 1-cm^2 pieces of copper, silver, and magnesium metal
- copper(II) nitrate solution
- silver nitrate solution
- microscope slides
- microscopes (preferred), hand lenses, or magnifying glasses

Procedure

Part 1: Formation Reaction: Anhydrous Copper(II) Sulfate + Water

1. Place about 2 g of anhydrous copper(II) sulfate in a test tube. Record the colour.
2. Hold the test tube by curling your fingers around it, so you can detect any temperature changes when you add water.
3. Using a dropper, add 5 drops of water to the anhydrous copper(II) sulfate in the test tube. Record any colour changes and temperature changes in your notebook.

Part 2: Formation Reaction: Iron + Oxygen

4. Light the Bunsen burner. Using the tongs, light one of the pieces of steel wool, and record the result.
5. Place about 1 g of manganese(IV) oxide in an Erlenmeyer flask. Add 5 mL of 3% hydrogen peroxide solution. Observe bubbles forming. This is pure oxygen gas.
6. Using the tongs and the Bunsen burner, light the second piece of steel wool and plunge it into the Erlenmeyer flask. Record the result.
7. Follow your teacher's instructions for disposing of all the chemicals you have used.

Part 3: Decomposition Reaction: Copper(II) Sulfate Pentahydrate

8. Place 5 g of copper(II) sulfate pentahydrate in a test tube.
9. Using tongs, heat the test tube over a Bunsen burner flame. Note particularly the formation of any product on the walls of the test tube at the top. Note any colour changes in the material at the bottom of the test tube. Record your observations.
10. Once the test tube is cool enough to touch, add a few drops of water to the product. Record any changes.

Part 4: Single Replacement Reaction: Metal Ion Solutions with Magnesium and Copper Metals

11. Use a microscope or magnifying glass for the following steps.
12. Put a piece of magnesium on a clean slide, and place the slide under the microscope lens. Focus on the edge of the magnesium.
13. Put a drop of silver nitrate solution on the magnesium. Record your observations. Note whether a reaction is occurring or not. If you are using a microscope, try to describe the growth of the metallic crystals.
14. Repeat steps 12 and 13 with the copper metal and the silver nitrate solution.
15. Repeat steps 12 and 13 using the silver metal with the copper(II) nitrate solution.
16. These solutions contain toxic metal ions. They require special disposal procedures. Follow your teacher's instructions for disposing of all the substances you have used.

Analyzing and Interpreting

Part 1: Formation Reaction: Anhydrous Copper(II) Sulfate + Water

1. What evidence is there that adding water to anhydrous copper(II) sulfate creates a chemical reaction and not just a mixture?
2. Is the reaction of anhydrous copper(II) sulfate with water exothermic or endothermic? Explain.

Part 2: Formation Reaction: Iron + Oxygen

3. Describe the differences between burning steel wool (a source of iron) in air and burning it in pure oxygen.
4. Rewrite the following sentence as a balanced formula equation. Include the state symbols.
   Solid iron reacts with oxygen gas to form solid iron(III) oxide.

Part 3: Decomposition Reaction: Copper(II) Sulfate Pentahydrate

5. What evidence is there that water is one of the products of the decomposition of copper(II) sulfate pentahydrate?
6. What is the other product of this decomposition reaction?

Part 4: Single Replacement Reaction: Metal Ion Solutions with Magnesium and Copper Metals

7. Do all the combinations of metals and solutions appear to react in the same way? Include the rate of the reaction and the shape of the crystals in your answer.
8. Translate the following sentences into balanced formula equations, showing state symbols:
   a) Magnesium metal reacts with silver nitrate solution to produce silver metal and magnesium nitrate solution.
   b) Copper(II) nitrate solution reacts with magnesium metal to give magnesium nitrate solution and copper metal.
   c) Silver nitrate solution reacts with copper metal to give silver metal and copper(II) nitrate solution.

Forming Conclusions

9. Write a summary describing the observations that you made that indicate a chemical reaction has occurred for each type of reaction.

Applying and Connecting

10. Drying agents are used to keep humidity low by absorbing moisture from the air. For example, hearing aids are often stored at night in an airtight container that has a drying agent inside it. The drying agent must be regenerated when it has reached its water absorption limit. Use the results of this investigation to suggest a means for producing a drying agent that indicates when it has reached its absorption limit.

Extending

11. Based on your results, which is a more reactive metal: copper or silver? Explain.

Double Replacement Reactions

Double replacement reactions commonly occur between two ionic compounds. Ionic compounds are always solids at room temperature, so these reactions happen in solution (that is, dissolved in a liquid), where the ions have the opportunity to mix. This type of reaction often results in the formation of at least one precipitate.

A general statement for a double replacement reaction is:

$AB + CD \longrightarrow AD + CB$

The ions in the first compound join with ions from the second compound. This is called a double replacement reaction because two new ionic compounds are formed.

There are several facts to keep in mind as this double replacement occurs:

- A and C are both positive ions. They will never pair up together because they repel each other.
- A and C will always appear first in formulas because positive ions are always written first.
- B and D are negative ions, so they will combine with any positive ions. They are always written second in formulas.

Example Problem A3.9

When aqueous lead(II) nitrate and aqueous sodium iodide are mixed, a bright yellow precipitate of solid lead(II) iodide forms (Figure A3.20). Another product also forms. It is aqueous. Write the balanced equation for this double replacement reaction.

FIGURE A3.20 The yellow precipitate lead(II) iodide was widely used as a paint pigment in the past. Lead-based paints were popular, and today lead paint dust is a health hazard in very old homes.

1. Write the word equation: The question does not name the second product. However, every ionic compound has two parts to its name. Look at the reactants and switch the parts of their names around.

   Word equation:
   lead(II) nitrate + sodium iodide ⟶ lead(II) iodide + sodium nitrate

2. Write the skeleton equation: All these substances are ionic. Use the solubility table (Table C) in Student Reference 12 to determine that PbI_2 is only slightly soluble. It appears in the equation as a solid.

   Skeleton equation: $Pb(NO_3)_{2(aq)} + NaI_{(aq)} \longrightarrow PbI_{2(s)} + NaNO_{3(aq)}$

3. Write the balanced equation: Remember—do not change any subscripts. Simply add coefficients.

   Balanced equation: $Pb(NO_3)_{2(aq)} + 2\ NaI_{(aq)} \longrightarrow PbI_{2(s)} + 2\ NaNO_{3(aq)}$

Practice Problem

9. Write the word, skeleton, and balanced equations for the following reactions.
   a) When aqueous copper(I) nitrate and aqueous potassium bromide are mixed, a precipitate of solid copper(I) bromide forms. Another product also forms.
   b) When aqueous aluminium chloride and aqueous sodium hydroxide are mixed, a precipitate of solid aluminium hydroxide forms. Another product also forms.

In Example Problem A3.9, $PbI_{2(s)}$ was one of the products. It is only slightly soluble. Another way to describe its solubility is to say that its solubility is low. Such substances do not dissolve very well. In chemical equations, they are shown as a solid *(s)* even if they are formed in the presence of water. The other product in the last example is $NaNO_{3(aq)}$. It is very soluble, and remains completely in solution unless most of the water is removed. Its solubility is high. When formed in water, it is shown in the chemical equation as aqueous *(aq)*.

Recall that the solubility table lists the combinations of ions that are slightly soluble and very soluble (see Table A2.13 on page 57 or Table C in Student Reference 12). Consult this table to determine the states of the products of replacement reactions.

Activity A11

QuickLab

Double Replacement Reactions

Purpose

To observe the formation of precipitates, and write the chemical equations that represent the reactions

Materials and Equipment

dropper bottles containing the following solutions:

- sodium iodide solution
- silver nitrate solution
- iron(III) chloride solution
- sodium hydroxide solution
- sodium carbonate solution
- calcium chloride solution

5 test tubes

test tube rack

Procedure

1. Using a separate test tube for each reaction, combine 5 drops of each of the following solutions. Check for the formation of a precipitate. If there is a precipitate, note its colour. If there is no precipitate, write NR (no reaction).

   a) sodium iodide solution and silver nitrate solution
   b) iron(III) chloride solution and sodium hydroxide solution
   c) sodium carbonate solution and calcium chloride solution
   d) sodium iodide solution and calcium chloride solution
   e) silver nitrate solution and sodium carbonate solution

2. Follow your teacher's instructions for disposing of all the substances you have used.

Questions

1. Which combinations above produced a precipitate? For each reaction that produced a precipitate, write a word equation and a skeleton equation. Then balance the equation.

2. Use the solubility table (Table C) in Student Reference 12 to determine the solubility of the products in each reaction. Which are only slightly soluble? Which are very soluble?

3. Write the state symbol for each compound in each equation. All the reactants are very soluble, so they are all aqueous *(aq)*. One product in each reaction is soluble, and the other is only slightly soluble. The slightly soluble product should be shown in the equation as solid *(s)*.

Predicting the Products of Chemical Reactions

Now that you have become familiar with five simple types of chemical reactions, you have the foundation for predicting the outcome of thousands of chemical reactions. Each reaction type follows a pattern that is different from the others. An examination of the reactants usually makes it possible to predict the identity of the products.

The first step in predicting what will happen in any reaction is to classify the reaction. Then, the products can be determined by following the pattern for that type. Table A3.3 summarizes the types of reactions.

TABLE A3.3 Summary of Types of Reactions

Type of Reaction	Reactants		Products
formation (synthesis)	A + B	$\longrightarrow$	AB
decomposition	AB	$\longrightarrow$	A + B
hydrocarbon combustion	$C_xH_y + O_2$	$\longrightarrow$	$CO_2 + H_2O$
single replacement	(A is a metal) A + BC (D is a non-metal) D + BC	$\longrightarrow$ $\longrightarrow$	B + AC C + BD
double replacement	AB + CD	$\longrightarrow$	AD + CB

Note that there are many kinds of reactions that do not fit into the five categories that you have just learned. There are even variations of these five types for which the products are difficult to predict.

FIGURE A3.21 Hexane ($C_6H_{14(l)}$), in part (c) of Example Problem A3.10, is one of the components of gasoline.

Example Problem A3.10

Classify each of the following reaction types. Predict the names of the product(s), and write the skeleton equation.

a) $PbBr_{4(s)} \longrightarrow$

b) $Ni(NO_3)_{3(aq)} + Na_2SO_{3(aq)} \longrightarrow$

c) $C_6H_{14(l)} + O_{2(g)} \longrightarrow$ (assume plenty of oxygen is available during the reaction)

d) copper + gold(III) chlorate $\longrightarrow$

e) zinc + sulfur $\longrightarrow$

a) $PbBr_{4(s)} \longrightarrow$

- Classify the reaction: Since there is only one reactant, this must be a decomposition reaction.
- Predict the names of the products: This compound contains the lead(IV) ion, Pb^{4+}, and the bromide ion, Br^-. It can be decomposed into the elements lead and bromine.
- Write the skeleton equation: Lead metal is shown by its element symbol in the products because it is an atom and has no charge. Bromine is diatomic.

Skeleton equation: $PbBr_{4(s)} \longrightarrow Pb_{(s)} + Br_{2(l)}$

b) $Ni(NO_3)_{3(aq)} + Na_2SO_{3(aq)} \longrightarrow$

- Classify the reaction: Two ionic compounds mean this is a double replacement reaction.
- Predict the names of the products: The nickel(III) ion, Ni^{3+}, will combine with the sulfite ion, $SO_3^{\,2-}$, to make nickel(III) sulfite. Similarly, the sodium ion, Na^+, will combine with the nitrate ion, $NO_3^{\,-}$, to make sodium nitrate.
- Write the skeleton equation:

Skeleton equation: $Ni(NO_3)_{3(aq)} + Na_2SO_{3(aq)} \longrightarrow Ni_2(SO_3)_{3(s)} + NaNO_{3(aq)}$

c) $C_6H_{14(l)} + O_{2(g)} \longrightarrow$

- Classify the reaction: The presence of a hydrocarbon with oxygen indicates that this is a hydrocarbon combustion reaction.
- Predict the names of the products: Assuming that there is plenty of oxygen available during combustion, the products will be carbon dioxide and water vapour.
- Write the skeleton equation:

Skeleton equation: $C_6H_{14(l)} + O_{2(g)} \longrightarrow CO_{2(g)} + H_2O_{(g)}$

d) copper + gold(III) chlorate $\longrightarrow$

- Classify the reaction: An element reacting with a compound is characteristic of a single replacement reaction.
- Predict the names of the products: The element is a metal, so it will replace the gold(III) ion, Au^{3+}. The gold is an element in the products, and since it is in the form of an atom, it will have no charge. The copper forms the Cu^{2+} ion. This is the most common form of the ion (2+), so it appears first in the periodic table. The Cu^{2+} ion exists in combination with the chlorate ion, $ClO_3^{\,-}$, to give the ionic compound copper(II) chlorate.
- Write the skeleton equation:

Skeleton equation: $Cu_{(s)} + Au(ClO_3)_{3(aq)} \longrightarrow Cu(ClO_3)_{2(aq)} + Au_{(s)}$

e) zinc + sulfur $\longrightarrow$ (assume that there is enough heat to start the reaction)

- Classify the reaction: Both zinc and sulfur are elements. This must be a formation reaction, with only one product.
- Predict the names of the products: Zinc is a metallic element and is shown in the reactants simply by its symbol, Zn. It has no charge in this form because it is an atom. Sulfur exists at room temperature as the molecule S_8, which is how it appears in the reactants. The product is the compound formed from the zinc ion, Zn^{2+}, and the sulfide ion, S^{2-}, to give zinc sulfide.
- Write the skeleton equation:

Skeleton equation: $Zn_{(s)} + S_{8(s)} \longrightarrow ZnS_{(s)}$

FIGURE A3.22 Sulfur and zinc combine in a formation reaction (part (e) of Example Problem A3.10).

Practice Problem

10. Classify each of the following reaction types. Predict the names of the product(s), and write the skeleton equation.
    a) $C_4H_{10(g)} + O_{2(g)} \longrightarrow$ (assume plenty of oxygen is available during the reaction)
    b) $Ca(NO_3)_{2(aq)} + Na_3PO_{4(aq)} \longrightarrow$
    c) calcium + silver nitrate $\longrightarrow$
    d) magnesium + oxygen $\longrightarrow$
    e) $AlCl_{3(s)} \longrightarrow$

Activity A12

QuickLab

Classifying Chemical Reactions

Purpose

To classify a variety of chemical reactions and write the chemical equations that represent them

Materials and Equipment

- aluminium foil
- copper(II) chloride solution
- cobalt(II) chloride hexahydrate
- silver nitrate solution
- calcium chloride solution
- sodium carbonate solution
- 10-cm length of copper wire
- Bunsen burner or hot plate
- 4 test tubes
- test tube rack
- 10-mL graduated cylinder
- tongs
- balance

Procedure

1. Roll up a 5 × 5 cm square of aluminium foil. Place 5 mL of copper(II) chloride solution in a test tube, and drop the aluminium foil into it. Record your observations.
2. Place 1 g of cobalt(II) chloride hexahydrate in a test tube. Use tongs to heat the test tube over a Bunsen burner. If you do not have a Bunsen burner, place the cobalt(II) chloride hexahydrate on aluminium foil and heat on a hot plate. Record your observations.
3. Place 5 drops of silver nitrate solution in a test tube and add the copper metal. Record your observations.
4. Place 5 drops of calcium chloride solution in a test tube. Add 5 drops of sodium carbonate solution. Record your observations.
5. These solutions contain toxic metal ions. They require special disposal procedures. Follow your teacher's instructions for disposing of all the substances you have used.

Questions

1. Write a word equation and a balanced formula equation for the reaction of copper(II) chloride solution with aluminium metal. Include state symbols.
2. Classify the reaction involving cobalt(II) chloride hexahydrate. Name one of the products of the reaction. What evidence was there that it was produced?
3. Classify the reaction of silver nitrate and copper metal. Only one of the products was a solid. What was it?
4. Classify the reaction involving calcium chloride and sodium carbonate. Write a word equation and a balanced formula equation for this reaction. Include state symbols. One of the products of the reaction is white. Indicate on the formula equation which one is white.
5. Which two reaction types that you studied did not occur in this experiment?

Example Problem A3.11

A solution of copper(II) nitrate is placed in an aluminium pot for storage. Almost immediately this proves to be a mistake. A sudden colour change and an increase in temperature indicate that a chemical reaction is taking place. Explain what is happening by using a word equation, a skeleton equation, and a balanced equation. Include state symbols in the balanced equation.

1. Identify the reactants and classify the reaction: An element and a compound are reacting together so this is a single replacement reaction. The element is a metal, so it will replace copper in the compound. Copper metal will be produced. The aluminium metal will form the Al^{3+} ion. This means the aluminium pot should look corroded.

2. Write the word equation:

   Word equation:
   aluminium + copper(II) nitrate $\longrightarrow$ copper + aluminium nitrate

3. Write the skeleton equation:

   Skeleton equation: $Al_{(s)} + Cu(NO_3)_{2(aq)} \longrightarrow Cu_{(s)} + Al(NO_3)_{3(aq)}$

   Both aluminium metal and copper metal are solids at or near room temperature. The copper(II) nitrate was in solution so it is aqueous. In the solubility table (Table C) in Student Reference 12, you can see that all nitrates are very soluble, so there will not be any precipitate from this new compound. It is aqueous.

4. Write the balanced equation:

   Balanced equation: $2\ Al_{(s)} + 3\ Cu(NO_3)_{2(aq)} \longrightarrow 3\ Cu_{(s)} + 2\ Al(NO_3)_{3(aq)}$

   We can conclude that the aluminium pot is corroding from the inside because of its reaction with the copper(II) nitrate solution. A pile of metallic copper powder should be sitting in the bottom of the pot. There is no pile of aluminium nitrate powder, because it is very soluble and remains in solution. Try plastic or glass next time!

*re*SEARCH

Most fertilizers contain ammonium (NH_4^+) salts and nitrate (NO_3^-) salts. These are synthesized from ammonia ($NH_{3(g)}$). Research the formation of ammonia from hydrogen and nitrogen using the Haber process. Begin your search at www.pearsoned.ca/school/science10

Practice Problems

11. A lead(IV) nitrate solution is placed in a zinc pot for storage. A sudden colour change and an increase in temperature indicate that a chemical reaction is taking place. Explain what is happening by using a word equation, a skeleton equation, and a balanced equation. Include state symbols in the balanced equation.

12. A reaction occurs when silver metal is placed in a solution of gold(III) nitrate. Write a balanced equation. Include the state symbols.